Name Key

Unit 15 Practice Test: Applications of Aqueous Equilibrium

1. Which of the following pairs of compounds could create a suitable buffer system?

   a. sodium sulfate and sodium hydroxide
   b. sodium chloride and hydrochloric acid
   c. carbonic acid and sodium chloride
   d. sodium acetate and acetic acid ← weak acid + conjugate base
   e. ammonia and potassium chromate

2. What is the pH of a solution that is 0.50 M in propanoic acid and 0.40 M in sodium propanoate. (K_a for propanoic acid = 1.3×10^{-5})

   a. -4.98
   b. 0.097
   c. 4.79
   d. 0.47
   e. 4.98

$$HP \rightleftharpoons H^+ + P^-$$
$$.50M - x \quad x \quad .40M + x$$

assume x is very small compared to concentrations

$$pH = pKa + \log\frac{(base)}{(acid)}$$
$$pH = 4.89 + \log\frac{(.40)}{(.50)} = 4.79$$

3. A buffer solution is prepared that is 0.50 M in propanoic acid and 0.40 M in sodium propanoate with a solution volume of 1.00 liters. (K_a for propanoic acid = 1.3×10^{-5}). What is the pH of the solution when 0.060 mol of NaOH(s) is added to the solution? Assume no change in solution volume.

   a. 4.91
   b. 4.77
   c. 4.67
   d. 3.97
   e. 4.97

	HP	P^-
I	.50	.40
C	-.06	+.06
E	.44	.46

$$pH = 4.89 + \log\frac{(.46)}{(.44)}$$
$$pH = 4.91$$

4. A chemist desires to create a buffer solution beginning with 1.00 liter of 0.200 M NH_3. How many moles of gaseous HCl must be introduced in order to produce a buffer of maximum capacity. Assume no increase in solution volume.

   a. 0.300 mol HCl
   b. 0.500 mol HCl
   c. 0.100 mol HCl
   d. 0.200 mol HCl
   e. None - the solution is already buffered

$$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$$

	NH_3	NH_4^+
I	.200	~0
C	-.100	+.100
E	.100	.100

Max. buffer when conc. of weak acid & conj. base are equal.

5. 200.0 mL of 0.200 M HCl is titrated with 0.050 M NaOH. What is the pH after the addition of 100.0 mL of the NaOH solution? Both strong

a. 0.82
b. 0.76
c. 0.93
d. 1.03
e. 1.45

$H^+ + OH^- \rightarrow H_2O$

	$[H^+]$	$[OH^-]$
I	.133	.017
C	-.017	-.017
E	.116	~0

$pH = -\log[.116] = .93$

Remember to calculate conc. after mixing. (diluted)

6. A 0.100M solution of NH_3 is prepared by dissolving 0.0050 moles of NH_3 into 50.0 mL of water. What is the concentration of NH_4^+ in solution when it reaches equilibrium? K_b for NH_3 is 1.81×10^{-5}

a. 0.00134M
b. 0.00425M
c. 1.81×10^{-5}M
d. 0.100M
e. 0.00500

$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$

	$[NH_3]$	$[NH_4^+]$
I	.100	0
C	-x	+x
E	.1-x	x

$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]}$

$1.81 \times 10^{-5} = \frac{x^2}{.1-x}$

assume x is very small

$x = 0.00134M$

7. The 50.0 mL of 0.100 M NH_3 from the previous problem is titrated with 0.025 M HCl. What is the pH of the solution after 100.0 mL of HCl has been added?

a. 4.74
b. 10.90
c. 8.99
d. 12.45
e. 9.15

	NH_3 (mol)	NH_4^+ (mol)
I	.0049	.000067
C	-.0025	+.0025
E	.0024	.0026

$pOH = pK_b + \log\frac{(acid)}{(base)}$

$pOH = 4.74 + \log\frac{(.0026)}{(.0024)}$

$pOH = 4.77$

$pH = 9.22$

8. K_a for HNO_2 is 4.5×10^{-4}. calculate the pH of a buffer solution made by mixing 0.225 mol of HNO_2 and 0.450 mol of $NaNO_2$ in enough water to make 1.00 liters of solution.

a. 3.65
b. 3.51
c. 2.98
d. 5.72
e. 3.90

$HNO_2 \rightleftharpoons H^+ + NO_2^-$

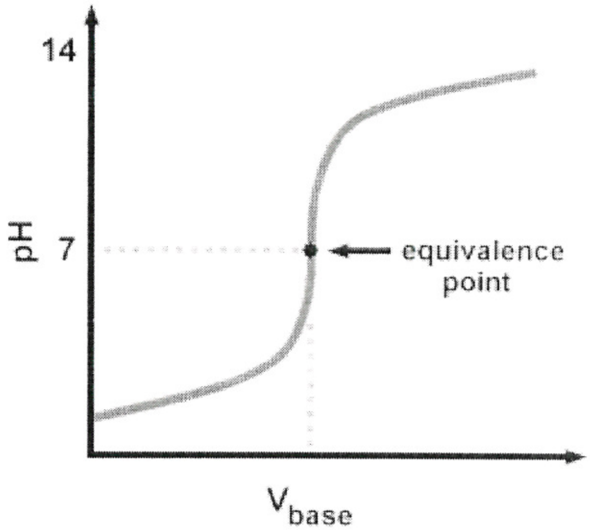
.225M x .450M

$pH = 3.347 + \log\frac{(.450)}{(.225)} = 3.65$

9. This image shows the titration of a:

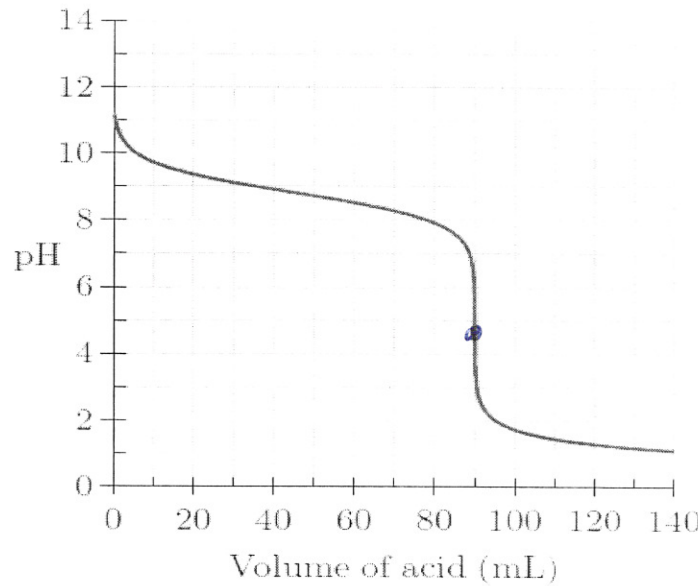
a. Weak acid with a strong base
b. Strong acid with a weak base
c. Strong acid with a strong base
d. Weak acid with a strong acid
e. Weak base with a strong base

Neutralizes at a pH of 7



10. This image shows the titration of a:
a. Weak acid with a strong base
b. Strong acid with a strong base
c. Strong base with a strong acid
d. Strong base with a weak acid
e. Weak base with a strong acid

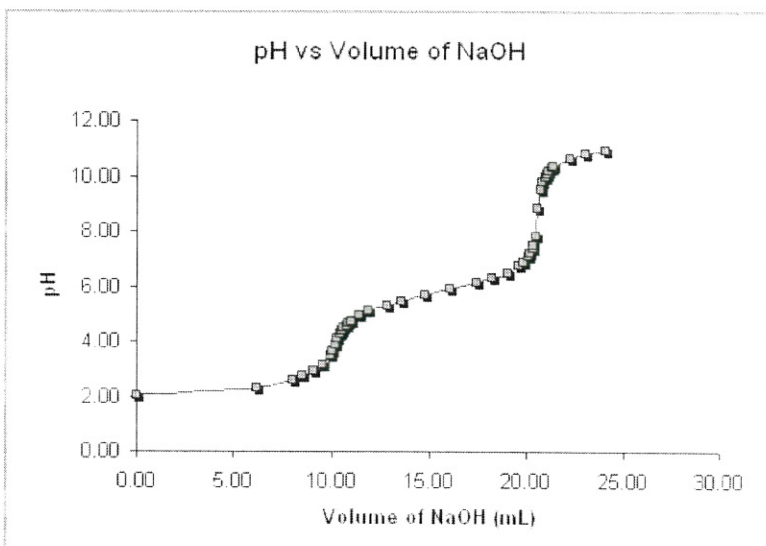
Neutralizes at a pH below 7.



11. An unknown acid is titrated with 0.100M NaOH producing the following titration curve. Which of the following acids are possible identities for the unknown?

a. $HClO_3$
b. H_2SO_4
c. H_3AsO_4
d. HF
e. $H_2C_2O_4$

Diprotic weak acid



12. The K_{sp} of $BaSO_4$ is 1.5×10^{-9}. What is the molar concentration of $Ba^{2+}(aq)$ in a saturated solution of $BaSO_4$?

a. 3.9×10^{-5} M
b. 6.7×10^{-15} M
c. 2.45×10^{-5} M
d. 2.25×10^{-18} M
e. 1.7×10^{-2} M

$BaSO_{4(s)} \rightleftharpoons Ba^{2+} + SO_4^{2-}$

$K_{sp} = [Ba^{2+}][SO_4^{2-}]$

$1.5 \times 10^{-9} = x^2$ $\quad x = 3.9 \times 10^{-5}$ M

13. Calculate the K_{sp} of $Ba(OH)_2$ given the fact that the solubility of $Ba(OH)_2$ in water is 0.107 mol/L.

a. 4.90×10^{-3}
b. 2.03×10^{-7}
c. 1.49×10^{-4}
d. 1.22×10^{-2}
e. 1.08×10^{-1}

$Ba(OH)_2 \rightleftharpoons \underset{x}{Ba^{2+}} + \underset{2x}{2OH^-}$

$K_{sp} = [x][2x]^2 = [.107][.214]^2$

$K_{sp} = .00490$

14. The above solution of Barium hydroxide is added to a solution of barium chloride. What will happen to the solubility of $Ba(OH)_2$?

a. solubility will increase
b. solubility will decrease
c. solubility will stay the same

common ion effect

15. 200.0 mL of 0.100M NaOH is added to 200.0 mL of 0.100M $AgNO_3$. If the Ksp for AgOH is 2.0×10^{-8}, will a precipitate be produced?

a. Yes
b. No
c. Impossible to tell

$AgOH_{(s)} \rightleftharpoons \underset{\substack{x \\ .05M}}{Ag^+} + \underset{\substack{x \\ .05M}}{OH^-}$

$Q = [.05][.05] = .0025$

$Q > K_{sp}$ so we have too many ions dissolved, thus it will precipitate.